CHM129

Acid-Base Equilibrium: Weak Acids and Bases (Practice)

1. A 0.100M solution of chloroacetic acid ($ClCH_2COOH$) is 11.0% ionized. Calculate $[H^+]$, $[ClCH_2COO^-]$, $[ClCH_2COOH]$ and K_a for chloroacetic acid.

$$ClCH_2COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + ClCH_2COO^-_{(aq)}$$

$$K_a = \frac{[H_3O^+][ClCH_2COO^-]}{[ClCH_2COOH]}$$

$$= \frac{(0.0110)(0.0110)}{(0.100 - 0.0110)}$$

$$K_a = 1.36 \times 10^{-3}$$

$$\% \text{ ion.} = \frac{[H^+]}{[HA]} \times 100$$

$$[H^+] = \frac{(11.0\%)(0.100\,M)}{100} = 0.0110\,M$$

	[HA]	$[H_3O^+]$	$[A^-]$
I	0.100	0	0
C	−0.0110	+0.0110	+0.0110
E	(0.100−0.0110)	0.0110	0.0110

2. The pKa of saccharin ($HNC_7H_4SO_3$) is 2.32 at 25°C. What is the pH of a 0.10M solution of saccharin?

$$HNC_7H_4SO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + NC_7H_4SO_3^-{}_{(aq)}$$

$$K_a = \frac{[H_3O^+][NC_7H_4SO_3^-]}{[HNC_7H_4SO_3]} = 4.8 \times 10^{-3}$$

$$\frac{(x)(x)}{0.10 - x} = 4.8 \times 10^{-3}$$

$$\underbrace{x^2}_{a} + \underbrace{4.8 \times 10^{-3}}_{b} x \underbrace{- 4.8 \times 10^{-4}}_{c} = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{-(4.8 \times 10^{-3}) \pm \sqrt{(4.8 \times 10^{-3})^2 - (4)(1)(-4.8 \times 10^{-4})}}{2(1)}$$

$x = \underline{0.015}$; $\cancel{-0.025}$ $\qquad [H_3O^+] = \underline{0.015\,M}$

$$pH = -\log(0.015) = \underline{1.81}$$

$$K_a = 10^{-pK_a}$$
$$= 10^{-2.32} = 4.8 \times 10^{-3}$$

	[HA]	$[H_3O^+]$	$[A^-]$
I	0.10	0	0
C	−x	+x	+x
E	0.10−x	x	x

$$\frac{[HA]}{K_a} = \frac{0.10}{4.8 \times 10^{-3}} = 20.9 < 400$$

Cannot assume x is small

3. What is the pH of a 0.15 M NH_3 solution? $K_b = 1.8 \times 10^{-5}$

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$

$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
I	0.15	0	0
C	$-x$	$+x$	$+x$
E	$0.15-x$	x	x

$\frac{(x)(x)}{0.15-x} = 1.8 \times 10^{-5}$

$\frac{[NH_3]}{K_b} = \frac{0.15}{1.8 \times 10^{-5}} = 8300 > 400$

Assume x is small

$\frac{x^2}{0.15} = 1.8 \times 10^{-5}$ Assume x is small

$x = \sqrt{(0.15)(1.8 \times 10^{-5})} = 1.6 \times 10^{-3} M = [OH^-]$

$pOH = -\log [OH^-] = -\log(1.6 \times 10^{-3}) = 2.78$

$pH = 14.00 - pOH = 14.00 - 2.78 = \underline{\underline{11.22}}$

4. Given that the Kb of ammonia (NH_3) is 1.8×10^{-5} and that for hydroxylammine (NH_2OH) is 1.1×10^{-8}, which is the stronger base? Predict which has the strongest conjugate acid. Determine Ka for NH_4^+ and NH_3OH^+. Was your prediction correct?

NH_3 has a larger K_b. NH_3 is stronger than NH_2OH.

Conjugate acids : NH_3 : NH_4^+

NH_2OH : NH_3OH^+

NH_3OH^+ is the stronger conjugate acid because it's the conjugate acid of the weaker base.

$K_{a, NH_4^+} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$

$K_{a, NH_3OH^+} = \frac{1.0 \times 10^{-14}}{1.1 \times 10^{-8}} = 9.1 \times 10^{-7}$ $\Leftarrow$ Stronger acid because its Ka value is larger